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THE AERATION TEST FOR SEWAGE EFFLUENTS.

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(*Read at the Meeting, April* 3, 1901.)

The value of periodical determinations of the changes in dissolved oxygen being now generally admitted, a uniform method of applying the test is under consideration. Many outdoor circumstances, such as space and time, soil, life, flow, and wind, cannot be imitated in the laboratory, so that the test so applied must always be to some extent an artificial one.

The agencies at work in the natural purification are: (1) Anaerobic and semi-anaerobic bacteria effecting hydrolytic or resolving changes without the presence of dissolved oxygen; (2) aerobic organisms, requiring it and consuming it in their life; (3) algæ and water-plants, living partially on the carbonic acid excreted by the bacteria and liberating oxygen required by the aerobic forms. This third class demands the presence of light, the other two work best in the dark, so that the cycle of changes depends on the action of the three classes, but cannot be well carried on in the same area. In the deeper layers of ponds and rivers there is very little free oxygen, and the changes are mainly anaerobic. In the upper layers oxygen is rapidly absorbed from the atmosphere. If the organic matter, however, be so large in quantity that the absorbed oxygen is not sufficient for the aerobic bacteria, it can be supplemented by the decomposition of nitrates, or by the oxygen disengaged by green algæ. A healthy condition would be that the balance of dissolved oxygen should be maintained, so that the three classes of organisms could work in their respective zones.

In examining effluents as to whether their quality is sufficiently good to allow them to be mixed with the local river water without putrescence, both the closed and the open tests are in use.

The former, introduced by Dr. Dupré in 1885, and now regularly worked by Mr. Scudder at Manchester, preserves the liquid in a full bottle, and determines the dissolved oxygen at the beginning and at intervals of one, two, three or more days.

This fermentation resembles that at the bottom of rivers and of deep ponds. But the dissolved oxygen in saturated water only amounts to about 1 part by weight in 100,000, so could not be sufficient to oxidize the organic matter of mixed liquids which, exposed to the air, would rapidly improve. I have just determined the behaviour of a town sewage as to dissolved oxygen under open and closed conditions, before and after passing through a septic tank.

	Dissolved Oxygen in c.c. per Litre.	
	Raw Sewage.	Tank Effluent.
Kept six days in closed full bottles	0·59	0·43
Exposed to air in a layer 1 foot deep :		
Twenty-four hours	1·81	2·06
Forty-eight hours	3·02	4·74

The odour in the closed bottle increased ; in the second case of exposure to air it was very slight after twenty-four hours, and none after forty-eight hours. Under natural conditions, therefore, the sewage would have become inoffensive.

More artificial is the system of conducting the closed process in the dark. Albert Levy suggested that in this way the action of bacteria could be discriminated from that of green algæ, which require light for evolving oxygen. Having ascertained the original free oxygen, he enclosed the water in a full bottle placed in the dark at 33° C. for forty-eight hours and determined the oxygen left ; the loss, divided by the original amount, was called the "coefficient of alterability," which he considered a measure of the quantity of bacteria. It, however, only regards a section even of the bacteria, and does not relate to natural conditions. The test is systematically carried out in Paris.

The Open Process.—The depth and breadth of bodies of water, the circulation of the fluid, the temperature, and the amount of light, have all an influence on the natural purification, and vary greatly at different places and times. The character of the organisms will also vary.

Trials made in the laboratory with small quantities give valuable information, but the conditions should be stated, and for purposes of comparison a standard procedure is needed. Dibdin and Thudichum (*J.S.C.I.*, June 30, 1900) allow a mixture of one part of effluent and one of fully aerated water to stand in an open vessel *of diameter equal to the depth of the liquid*, and determine the dissolved oxygen daily for a certain number of days ; it should not fall below 50 per cent. of the maximum possible at the temperature. "Whenever the aeration of the water of the Thames was found to exceed this minimum there was no suggestion of nuisance, and fish appeared in the areas within which they could not live during other periods."

In regard to the "fish test," Duncan and F. Hoppe-Seyler (*Zeit. Physiol. Chem.*, xvii., 165) found that perch, trout, etc., could thrive in water holding 3 to 4 c.c. of oxygen in solution per litre, but soon died when only 1·7 c.c. was present. With reference to the absorption of oxygen by water, the same observers (*Zeit. Physiol. Chem.*, xvii., 147) filled tubes of different lengths with boiled water and exposed them to the atmosphere for varying periods, noting the temperature and barometric pressure. From the results they calculated tables of the diffusion of the gases into the water at different depths below the surface. The total amount of gas absorbed

increased with the duration of exposure, but the daily increment in the amount absorbed gradually diminished.

Mr. Dibdin also remarked (*loc. cit.*) : "If the aeration fell below 50 per cent., it showed that a quantity of organic matter was present on which a number of aerobic organisms were being maintained, with the result that they were absorbing oxygen quicker than it could be reabsorbed by the water from the atmosphere. But if 50 per cent. or over of aeration were steadily maintained, and if the percentage steadily arose, it was clear that the sample was progressing towards absolute purification." From the results of London County Council experiments, he concluded that oxygen-free water when suddenly exposed to the air took up during the first hour about 15 per cent. of the maximum absorption, and that as the quantity of oxygen absorbed in the water increased the rate of absorption decreased, "so that at some point a balance would be reached." But sewage, although nearly free from oxygen, is not comparable with boiled distilled water. When the oxygen as it dissolves is taken up by substances in the water, the rate of absorption will remain rapid. To imitate approximately the natural circumstances, I have mixed boiled water with ammonio-cuprous chloride and exposed it to air under various conditions, judging the absorption of oxygen by the tint of blue produced.

I. Three tubes of equal diameter ($\frac{1}{2}$ inch), and respectively 18, 12 and 6 inches long. After twenty hours the colours were inversely as the lengths. In sixty hours they were nearly equal, and all about half saturated. After eighty-four hours they were equal, but still short of saturation.

II. The amount of cuprous salt was made equivalent to the maximum dissolved oxygen at the temperature. The tubes were 12 inches long and 1 inch wide. The first tube gave :

Hours	1	2	3	4	5	22
Oxygen, c.c. per litre	1·0	1·5	2·5	3·0	3·0	saturated

A second tube, probably owing to more constant conditions, took a longer time to reach the maximum. It gave in one hour, 1·7 c.c. oxygen ; two hours, 2·6 ; three hours, 3·0. After this the increase of colour for some hours was so slow as not to be measurable. In eighteen hours it was equal to 3·5 c.c. oxygen ; the upper 7 inches had nearly reached saturation blue, the lower 5 were almost colourless. Twenty-six hours, 4 c.c. oxygen ; thirty-eight hours, colour 9 inches down, oxygen 5 c.c. Saturation reached in fifty hours.

III. Same as last, but with wider vessels, depth and width 3 inches.

Hours	1	2	3	4	20
Oxygen, c.c. per litre	1·5	2·5	2·7	3·0	saturated

In moderately shallow layers the results seem to depend less upon depth and width and amount of oxidizable matter than upon convection currects and vibration of the liquid. In the next series of experiments actual sewages were mixed with tap-water in equal proportions, placed in cylinders of varying sizes, and exposed to air and diffused light at ordinary temperatures. The liquids used were :

A. In experiments 1 and 2, a town effluent from sewage works, containing 4 parts of chlorine per 100,000.

B. Experiments 3, 4 and 5, a town sewage with 10 parts Cl.

C. Experiments 6 to 10, a foul-smelling sewage with dark earthy sediment; chlorine, 5·5 parts.

The physical conditions were:

No.	Depth, Inches.	Width, Inches.	Surface Area, Square Inches.	Volume.		Relative Volumes, Approximate.	Surface to Volume, Approximate.
				Cubic Inches.	C.c.		
1 and 6	6	2	3·14	18·85	309	1	1 : 6
2 and 7	12	2	3·14	37·7	618	2	1 : 12
3 and 8	6	5	19·6	117·6	1,929	6	1 : 6
4 and 9	6	$2\frac{3}{4}$	5·9	35·4	580	2	1 : 6
5 and 10	2	10	78·6	157·2	2,578	8	1 : 2

Dissolved oxygens:

No.	C.c. per Litre.			Percentage of Saturation.		
	24 Hours.	60 Hours.	120 Hours.	24 Hours.	60 Hours.	120 Hours.
A. 1.	5·49	5·35	—	89	81	
A. 2.	4·8	5·2	—	77·5	79	
B. 3.	3·3	6·1	—	53	92·5	
B. 4.	3·35	6·05	—	54	92	
B. 5.	4·2	6·25	—	68	95	
Tap-water	6·2	6·6	—	100	100	
		72 hours			72 hours	
C. 6.	2·25	3·13	—	35·4	52	
C. 7.	1·46	2·36	—	23	39	
C. 8.	3·0	4·20	5·64	47	70	97
C. 9.	2·68	4·08	—	42	67	
C. 10.	3·95	5·25	5·58	62	87	96
Tap-water	6·36	6·02	5·8	100	100	100

After sixty hours Nos. 1 and 2 were nearly colourless, and had scarcely any odour; trace of nitrite; nitrate = 1·4 parts N per 100,000.

Nos. 3, 4, and 5 after sixty hours showed a yellowish tint, with floating mycelium or zoogloœa, a distinct urinous odour, a heavy trace of nitrite, and no nitrate.

In series C, Nos. 6 to 10, after twenty-four hours, the settled liquid was opalescent, had a slight brownish tint and very little odour; nitrite, 0·15; nitrate, 0·368 part per 100,000.

After seventy-two hours there was no odour; the liquid was clear and nearly colourless. Nitrite, 0·05; nitrate, 0·42.

After 120 hours No. 8 contained 5·64; No. 10, 5·58; tap-water, 5·8; so that both were practically saturated. They were clear, with no odour and only a faint colour.

Conclusions.

1. The width of the sample has by itself very little influence on the aeration, provided the proportion of width to depth be great enough to allow sufficient circulation. A width of 2 inches, with a depth of 12 inches, is too narrow to satisfy this condition.

2. The depth, as would be expected, has a great effect; the amount of dissolved oxygen in large volumes of water diminishes with the distance from the surface, unless circulation interferes. Thus, in Adeney's tables the water of the reservoir contained at 5 feet depth 5·71 c.c. of O per litre, at 20 feet only 2·00, due to the action of anaerobes.

3. Too great a depth and too shallow a layer should be avoided, and the relation between surface and volume should be uniform in different experiments. In Mr. Dibdin's proposal of equal widths and depths (*J.S.C.I.*, June, 1900, p. 498), the ratio of surface to volume varies in each case.

4. The best directions for carrying out this form of the test uniformly seem to be: The samples should be kept at a uniform temperature in cylinders 6 inches deep and 3 inches wide, exposed to the air and light, but protected from dust. The volume will be about 600 c.c., and the ratio of surface to volume 1 to 6.

5. Even with uniform physical conditions, the action varies greatly with the nature of the organisms in the sewage and *in the water used for dilution.* Water from a water-supply that has been filtered is far less active than river water.

Influence of Organisms on the Absorption.

Three samples of liquid—

A. Water vigorously boiled in a flask for two hours;

B. Sewage similarly boiled, with additions of boiled water to maintain the volume;

C. Raw sewage previously kept in a closed vessel—

were exposed to air in a cool, quiet room in beakers, depth and width of liquid 3 inches, and the dissolved oxygen determined at intervals by the modified Winkler's process. As B and C contained nitrites and organic matter, previous permanganation was used in the manner I have described in a recent paper.*

No.	Hours.	Temp. C.	A. Boiled Water.			B. Boiled Sewage.			C. Raw Sewage.		
			C.c. O.	Roscoe's Maximum.	% of Saturation.	C.c. O.	Roscoe's Maximum.	% of Saturation.	C.c. O.	Roscoe's Maximum.	% of Saturation.
1	1	14·0	3·17	7·12	44·5	3·12	7·12	43·8	3·16	7·12	44·4
2	2	14·5	3·81	7·04	54·1	3·23	7·04	45·9	3·82	7·04	54·2
3	3½	15·5	4·41	6·89	64·0	4·12	6·89	59·8	4·26	6·89	61·8
4	5	14·0	4·84	7·12	68·0	4·70	7·12	66·0	4·42	7·12	62·0
5	11	14·5	5·69	7·04	80·8	5·13	7·04	72·8	4·05	7·04	57·5
6	25	14·5	6·71	7·04	95·3	6·22	7·04	88·3	3·0	7·04	42·6

* Analyst xxvi., p. 141.

The boiled sewage has behaved nearly in the same way as ordinary water. In the raw sewage the absorption is at first normal, but after five hours the consumption of oxygen by fermentations overpowers the absorption from the air, and the amount in solution sinks, till in twenty-five hours it is considerably below 50 per cent., so that by this datum the liquid would be condemned as putrescible.

ON ALKALINE WATERS FROM THE CHALK.

By W. W. Fisher, M.A.

(*Read at the Meeting, May* 1, 1901.)

The standards of purity for natural waters are necessarily based upon the ascertained composition of unpolluted samples, which will exhibit various differences according to their origin and the geological formations from which they are obtained. The extent to which waters are rendered as "hard" or "soft" depends roughly upon the quantities of calcium or magnesium compounds taken up by them, and we expect water-supplies from limestones or chalk to be hard, while those from sandstones (with some exceptions), shales, and igneous rocks, are as a rule soft in character.

Attention has, however, been repeatedly drawn to the fact that waters from deep wells in chalk in London and other places where the chalk is covered by the London clay and tertiary beds are entirely different in character from the waters obtained either from natural springs or deep wells in chalk districts where there is no such covering of impervious deposits.

The composition of a few typical chalk waters of the ordinary kind is shown in Table I., from which it appears that such waters contain from 20 to 30 grains per gallon of dissolved solid constituents. These are chiefly carbonate of lime, with a little magnesia, with small quantities of alkaline chlorides and sulphates. These waters are always hard to an extent proportional to the lime and magnesia in solution.

On the other hand, the waters obtained from the chalk beneath the London clay contain only mere traces of lime and magnesia as carbonates, while considerable quantities of alkaline salts, mainly consisting of sodium chloride, sulphate, and carbonate, are present, and these waters, consequently, are soft and alkaline.

In Table II. are given some examples of waters of this type derived from various localities, ranging from near Reading in Berkshire to Chelmsford in Essex, and in some of these waters the excess of saline constituents is very marked. The Essex sample contains 35·7 grains per gallon of common salt, the total solids being 74·8, with 7·3 grains of anhydrous sodium sulphate and a substantial amount of sodium carbonate; while the Wokingham water with 66 grains of total solids, contains 41·2 grains of common salt, with sodium sulphate and carbonate in distinct but smaller proportion. The water from Clewer, near Windsor, is similar in character. Many of the waters, it will be noticed, contain traces of free ammonia, but the proportion of nitrates is always small.

No.	Locality.	Depth in Feet.	Total Solids.	Chlorine in Chlorides.	Nitrogen in Nitrates.	Ammonia.	Albuminoid Ammonia.	Oxygen absorbed in 3 Hours.	Hardness.	Remarks.
3188	Slough, Bucks	211	22·68	1·0	0·210	0·025	0·005	0		
3055	Lockinge, Berks	60	33·0	1·7	0·154	0·002	0·007	0·035		
2217	Fawley, Bucks	330	38·0	2·1	0·84	0·004	0·003	0	26·2	
2747	Henley-on-Thames, Oxon ...	280	26·04	1·1	0·329	0·003	0·003	0·003	18·0	
2473	,, ,, ,, ...	100	25·2	1·0	0·301	0·003	0·004	0·005		
2777	Devizes, Wilts, well on Downs	(?)	19·04	0·8	0·035	0·003	0·003	0·010	16·0	
2808	Beaconsfield, Bucks ...	270	22·4	1·0	0·210	0·001	0·001	0		
2817	Great Missenden, Bucks ...	(?)	22·1	0·77	0·210	0·003	0·001	0·007		
3025	Amersham, Bucks, Waterworks well	150	19·88	0·9	0·091	0·003	0·001	0·002		
1525	Near Goodwood, Sussex ...	55	20·4	1·5	0·147	0·003	0·001	0·007	16·7	
1350	Savernake, Wilts	(?)	26·0	1·3	0·175	0·012	0·001	0·015		
2532	Hungerford, Berks	(?)	24·9	1·0	0·315	0·003	0·006	0·007	17·5	
2397	Wantage, Berks, from Downs	—	22·4	0·8	0·154	0·003	0·005	0·008		
2198	Farnham Royal, Bucks ...	60	19·0	1·8	0·231	0·001	0·003	0·023		
2166	Near Princes Risboro', Bucks	(?)	18·5	0·77	0·357	0	0·003	0		
2543	Great Marlow, Bucks ...	260	23·2	1·0	0·259	0·001	0·001	0		

TABLE II.—WATERS FROM CHALK BELOW LONDON CLAY. (*Analysis in Grains per Gallon.*)

No.	Locality.	Depth in Feet.	Total Solids.	Chlorine in Chlorides.	Nitrogen in Nitrates.	Ammonia.	Albuminoid Ammonia.	Oxygen absorbed in 3 Hours.	Hardness.	Remarks.
1539	Faversham, Kent ...	74	40·0	4·9	0·05	0·006	0·004	0·015	29·2	
2057	Aldermaston, Berks ..	260	23·0	1·1	0 04	0·008	0·003	0	13·7	Slight alkalinity. 9·1 after boiling.
1888	Peckham	259	35·0	2·6	0·014	0·021	0·004	0·005	29·6	
1878	Chelmsford, Essex ...	200	74·8	21·7	0·07	0·003	0·002	0·042	2·5	Na_2SO_4 7·3. Alkalinity 29·1, as Na_2CO_3.
1644	Chelsea	278	50·0	11·2	0·05	0·035	0·005	0·010	6·2	Alkaline.
2407	Wokingham, Berks...	408	66·0	25·1	0·05	0·003	0·003	0·003	—	Na_2SO_4. Alkalinity 18·9, as Na_2CO_3.
2261	Datchet, Berks ...	170	25·0	1·9	0·04	0·004	0·003	0·010	8·2	Alkaline.
2083	Clewer, Berks ...	300	62·0	17·9	0·05	0·039	0·003	0·019	11·5	Alkalinity 18·5, as Na_2CO_3.
3194	Hook, Hants ...	600	38·3	7·0	0·035	0·067	0 003	0	17·5	Alkaline.
2458	Wallingford, Oxon ...	(?)	37·0	2·2	0·014	0 028	0·004	0·010	—	Alkaline.

Numerous wells have been sunk into the chalk below London, and careful analyses have been made in some cases of the waters obtained, and on this account the Government wells near Trafalgar Square, another at the Royal Mint, and the deep well at Guy's Hospital, are of special interest as affording an opportunity of learning whether any changes in salinity have taken place over a period of more than fifty years. After application to H.M. Office of Works, by the kindness of Mr. Westcott, Engineer-in-charge, I was able to get a sample of the water, which on analysis yielded the results given in Table III., in which also the results of former analyses are embodied.

From these results the approximate quantities of the dissolved constituents are calculated as follows :

TRAFALGAR SQUARE WELL-WATER.

Composition in Parts per 100,000.

Total solids by evaporation dried at 100° C.	85·7
Calcium carbonate	4·29
Magnesium carbonate	3·15
Sodium carbonate (anhydrous)	18·82
Sodium sulphate (anhydrous)	28·64
Sodium chloride	27·34
Potassium chloride	2·16
Ferric oxide and phosphate	0·40
Silica	1·0
Total	85·80

A comparison of the water in 1900 with the earlier samples of 1846 and 1848 shows, as regards the total saline constituents, a diminution, but since 1857 there is no substantial change in the amount—during forty-four years. It will be seen, also, that the chlorine, sulphuric acid, lime and magnesia have undergone no important alteration for upwards of fifty years ; there is a striking change in the potash between 1848 and 1857, but since that time the alkalies are practically constant in quantity.

These facts are very significant, and of great interest as indicating the permanent character of the conditions affecting the supply. Although, as will be shown, the level of the water is slowly falling, from the water being pumped out from the chalk at a rate in excess of that coming in by percolation, it might have been expected that the salinity would become less ; but, judging from our present knowledge, the water will retain its characteristics for an indefinite period.

Several attempts have been made to explain the abnormal characters of the waters under consideration, the most popular, perhaps, being the suggestion of infiltration of sea-water (see Braithwaite, Proceedings Institute Civil Engineers, xiv., p. 507). The hypothesis, however, is clearly untenable, and does not account for the known facts. If such infiltration were possible in Essex, it hardly could be so in London, and still less at Wokingham. Further, the chalk is full of water under greater hydrostatic pressure from the hills north and south than from the bare chalk at Erith in the tidal river, and, so far from sea-water entering there, it is well known that strong springs of fresh water discharge from the chalk into the river in that

TABLE III.—ANALYSES OF TRAFALGAR SQUARE WELL-WATERS AT DIFFERENT TIMES. (*In Parts per* 100,000.)

Date.	Locality.	Total Solids.	Ch'orine.	SO_3.	CO_2.	SiO_2.	Fe_2O_3. P_2O_5.	CaO.	MgO.	Na_2O.	K_2O.
1846	Brande	94·4	22·3	15·9	10·6	1·0	—	2·52	1·66	(54·0	)
1848	Abel and Rown y ... (Q. J. Chem. Soc. I. 97)	99·15	17·39	16·02	14·6	1·3	0·99 0·23	2·61	1·56	35·9	10·55
1857	Campbell ... ... (Q. J. Chem. Soc. IX. 22)	84·97	16·51	17·19	8·3	0·57	0·97	2·18	1·40	39·7	1·02
1869	Frankland ... ... (Riv. Poll. Com. VI. Report)	83·40	16·55	—	—	—	—	[5·9 =	Hard-ness]		
1900	W. W. Fisher ...	85·7	17·6	16·14	11·35	1·0	0·40	2·40	1·50	38·0	1·35

TABLE IV.—ALKALINE WATERS FROM LIMESTONES UNDER CLAYS. (*Analysis in Grains per Gallon.*)

No.	Locality.	Depth in feet.	Total Solids.	Chlorine in Chlorides.	Nitrogen in Nitrates.	Ammonia.	Albu-minoid Ammonia.	Oxygen absorbed in 3 Hours.	Hard-ness.	Remarks.
2749	Newington, Oxon, below Gault Clay..	85	406	148·5	0·014	0·072	0·003	0·018	—	Alkaline = Na_2CO_3 17·5; SO_3 = 88 2 ; CaO trace.
3008	Warborough, Oxon, below Gault Clay...	177	545	190·7	0·049	0·252	0·006	0·044	—	Alkalinity = 12·5 Na_2CO_3; SO_3 = 106 9; CaO trace.
2991	Bourton, Berks, below Kimmeridge Clay (1899) ...	265	152	40·5	0 049	0·050	0·006	0·031	2·8	Alkalinity = 39·2 Na_2CO_3; SO_3 abundant ; CaO trace.
2339	Burcote, Oxon, below Kimmeridge Clay...	190	52	5·8	0·028	0·040	0·001	0·020	1·4	Alkalinity = 21·9 Na_2CO_3.
3027	Near Bicester, Oxon, below Oxford Clay	spring	47·6	4·2	0·070	0 007	0·009	0·470	—	Alkalinity = 31·5 Na_2CO_3; contains peaty organic matter and sodium sulphate ; CaO trace.
2528	Oxford, below Oxford Clay	250	672	173·0	0	0·101	0·045	0·394	—	Alkalinity = 28·0 Na_2CO_3; SO_3 = 201 ; CaO = 20 ; Na_2SO_4 = 350.
	Northampton, below Upper Lias Clay ...	200	61	9·9	—	0·050	—	—	—	Mr. A. C. Chapman's analysis: Na_2CO_3 = 21·70 ; NaCl = 16·14; Na_2SO_4 = 18·10; $(CaMg)CO_3$ = 3·85.

locality. On chemical grounds, too, it is clear that no mixture of sea-water with normal chalk-water could resemble our saline chalk supplies, since the relative amounts of chlorides and sulphates are entirely different, and such a mixture could not contain free alkaline carbonate. Further, there is no relation between the amount of salinity of any waters of this class and the distance of the well from salt-water or the sea.

Dugald Campbell also argued against the sea-water theory (*Amer. Journ. Chem. Soc.*, ix., p. 22), and came to the conclusion that the origin of the water is not the ordinary chalk stratum, but it comes from a separate and independent source, and suggests the tertiary beds above as the source of the saline matters.

In an account of Guy's Hospital well, Professor Odling (1860), who discussed this question, favoured the explanation that the most probable source of the carbonated alkali is to be found in the decomposition of alkaline silicates, chiefly silicate of soda existing in the strata above the chalk. Indeed, he showed that readily decomposed silicate of soda was an actual constituent of the sand stratum overlying the chalk obtained from the borings at Guy's Hospital well. But he makes also the important point that in a well at Vauxhall the chalk-water contains more carbonate of soda than the water from the sand above, and less sulphate of soda than the latter. There is much to be said in favour of this view, but it is not entirely conclusive, since the exposed tertiary beds do not yield waters of an alkaline or saline character, but only those, like the chalk, buried beneath the London clay, and subject to similar conditions.

In looking for a satisfactory explanation of this question, I have been strongly impressed by the fact that the alkaline waters are not peculiar to wells sunk through London clay, but are commonly (almost universally, I may say) met with in wells which, in the counties of Oxford, Berks, or Bucks, are sunk through clay formations into the porous limestone or other beds below. Many such instances are met with in wells sunk through gault, or Kimmeridge, or Oxford clays, and I hardly know an instance to the contrary. One or two examples have come to my knowledge of wells sunk into the porous middle lias beds between the upper and lower lias clays which yield saline and alkaline waters.

In Table IV. I have given a few examples of such waters.

We are therefore led to the conclusion that the origin of the saline matters must be sought for in the beds themselves, and that the overlying impervious clay is a necessary contributing factor in the case.

Now, in the Thames Basin London clay is a very extensive thick deposit, which, from Hungerford in Berks to the east coast of Essex, extends over a distance of 120 miles. The width in London is about thirty miles from north to south, and in Essex rather more. The north of Kent is partly covered by a long stretch of this deposit. Obviously, no water can pass downwards through this impervious layer, and the water which reaches the chalk below must enter beneath the edges of the clay, and the supply to the Trafalgar Square well comes either from the Hertfordshire chalk downs, a distance of fifteen miles underground, or those near Croydon, a distance of about nine or ten miles. Below the clay the chalk and tertiary beds are fully charged with water, and at the edge of the clay strong springs, fed by the surplus waters, break out; such waters being of the normal type of chalk supplies. Owing to the depression of the chalk in the centre, the waters have no

natural escape open to them, and would remain indefinitely in contact with it, and even when, owing to wells being sunk, a portion is removed by pumping, it is clear that a long time is required for the underground water to travel ten or fifteen miles. The chalk itself being remarkably compact, partly owing to the pressure from above, and more perhaps to the absence of such fissures and channels as would be produced by the ordinary action of flowing water, great resistance is offered to the passage of water through it. All the conditions favour the taking up of any soluble salts by water in contact with such chalk. It may also be pointed out that uncovered chalk, being subject to percolation for ages past, has lost the greater part of the easily-dissolved constituents, while the buried chalk, having no drainage outlets, has not been washed in the same way, and the soluble matters to a large extent remain.

These considerations appear sufficient to explain the increased proportion of chlorides and sulphates in such waters, and I hope to bring before the Society evidence of the progressive increase in saline constituents after an increasing length of underground passage of water through cornbrash under Oxford clay. The explanation of the presence of sodium carbonate, although a little more complicated, is essentially the same in principle. I have observed that as a fact nearly all chalk and limestone waters contain minute quantities of alkaline carbonates, which can be detected by the action on turmeric paper, and further traces of sodium carbonate are frequently present in chalk or limestone. In the underground journey the proportion of alkaline carbonate will gradually increase, and as it increases the calcium salts will diminish, until mere traces only are left. Neither chloride nor sulphate of calcium or magnesium will remain, and only limited quantities of carbonates. The amount of calcium carbonate in solution will ultimately depend on the amounts of carbonic acid in the water available as a solvent, so that some of the alkaline chalk waters are harder than others for this reason. In support of this view, I have examined specimens of chalk from a boring made some years ago in the Tottenham Court Road, one from a depth of 500 feet, the other from 800 feet. The cores are in the Geological Department of the University Museum, and I am indebted to the kindness of Professor Sollas for some fragments.

Both specimens contained distinct traces of alkaline carbonate, sulphate and chloride, and when extracted with redistilled water in a platinum vessel gave distinct reactions with turmeric paper. After evaporation the residue was redissolved and filtered, and the amount of sodium carbonate estimated by titration with decinormal acid was 0·03 per cent. of the weight of the chalk. The chalk from 800 feet deep was sufficiently alkaline to affect moist turmeric paper when a little of the powder was placed upon it.

An analysis of the sample of chalk from a depth of 500 feet gave the following results :

Calcium as carbonate	93·90
Silica (including soluble silica)	4·20
Iron oxide (with P_2O_5)	0·50
Magnesium carbonate	0·70
Sodium sulphate	0·28
Sodium chloride	0·20
Sodium carbonate	0·03
Water and loss	0·19
	100·00

Note.—It has been shown by Way (*Journ. Royal Agric. Soc.*, vol. xii., p. 544) that the lower chalk contains considerable amounts of silica and traces of alkaline salts.

Another specimen from a depth of 800 feet yielded alkaline carbonate in the proportion of 0·05 per cent.

The level of the water in the chalk wells below London is gradually sinking, and the Royal Commission on Metropolitan Water-supply, 1893, collected information as to the condition of many such wells (Report, Appendix, C. 4, p. 157). The depth of the Trafalgar Square well was then given as 384 feet, of which 241 feet was above the chalk, and 143 feet in the chalk. The water-level in October, 1900, was 174 feet from the surface, which is still 67 feet above the top of the chalk. The levels at various dates, according to the above report, which was prepared by Mr. A. R. Binnie, the Engineer to the London County Council, are as under: 1847, – 58·1 below Ordnance datum; in 1878, – 83·35; in 1888, – 100·68; in 1891, – 108·7; in 1900, – 117·1. Mr. Westcott, the Engineer-in-charge, informs me that the water does not appear to fall so fast recently, but it would appear that the level sinks about 1 foot per annum. The water clearly cannot filter or travel through the chalk for so great a distance as fast as it is being removed by pumping, and in the future possibly the level will be still further reduced.

Discussion.

Mr. Dibdin said that the question of variations in the quality of water obtained from certain deep wells in London had been one which had exercised the minds of all water analysts, who had hitherto generally been content to accept the explanation of Mr. Dugald Campbell—that the alkalinity of these waters was due to the solution of salts in the Thanet sands. But, if that explanation were correct, he (Mr. Dibdin) could not understand why the solution should not be much greater than was generally the case. He thought that the explanation suggested by Mr. Fisher, namely, the gradual solution of the soluble salts in the chalk during the progress of the water in its filtration through the chalk from the outside areas, was altogether far more reasonable. He would have been glad if Mr. Fisher had been able to give a few further analyses of waters drawing their supply from the Thanet sand. Any moderate variations that might be found in such waters would not in the slightest degree vitiate Mr. Fisher's argument, but would show how widespread was the variation in the quantity and character of their mineral constituents. There was another well close to that at Trafalgar Square, namely, in St. James's Park, of almost identical character; and on the south side of the Thames there were several wells yielding supplies of a similar character; while further down the Thames waters occurred which appeared to be mixtures of these soft alkaline waters with the hard water.

Mr. Richmond said that he had recently examined a sample of water from Bourton, near Swindon, which he believed to be from practically the same source as the one from that neighbourhood referred to by Mr. Fisher. There was a supply of a similar character near Shrivenham Station, but in that case the well was only a few feet deep; and on the other side of the Great Western Railway, about a mile and a half away, where the ground rose slightly, he had found the same water from

a well between 50 and 60 feet deep. On the other side of Swindon, three or four miles away, where the ground rose again, practically the same water occurred, the well in this case being about 100 feet deep. In all cases there was a large quantity of sodium sulphate, and the water was always alkaline. He had found, however, in going from Bourton to Swindon that there was a slight change in the water, the chlorine becoming less and less as one went further west. The chlorine in the Bourton water was about the same as that in Mr. Fisher's sample ; but on the other side of Swindon, although the sulphuric acid kept fairly constant, the chlorine fell to about 22 grains per gallon.

Mr. W. T. Burgess said that in some cases the alkalinity as determined by direct titration would include a certain quantity of carbonate of lime. In order to estimate the alkalinity of water, it had been his practice for a long time, in the case of waters containing carbonate of soda, to evaporate to dryness, extract with boiling water, and determine the alkalinity of the filtrate.

Mr. Blount said that, before one could accept so revolutionary a hypothesis as that which had been put forward as to the cause of the alkalinity of these waters, one must be quite sure of the data upon which it was founded, and he would suggest that Mr. Fisher would add greatly to the value of the paper if he would state with precision the exact methods of analysis he preferred to employ. In particular, he would ask that the method used for determining the alkalinity should be specifically stated, and also the data on which the author had been led to assume that a given water contained so much carbonate of any sort. In his own practice in water analysis, whenever a doubtful case occurred, he found that any titration method was to be put aside, the only satisfactory method in such cases being to evaporate nearly to dryness and determine the carbonic acid gravimetrically, as one would do in the case of chalk. The assumption that the alkalinity, however carefully determined, really represented the carbonic acid present in the water, and was a proper basis on which to calculate the carbonates of the mineral constituents, was apt, he thought, to lead to erroneous conclusions. The determination of each mineral constituent by some method that was quite irrefutable was the only means, not merely of arriving at a correct analysis of a water, but also of founding satisfactorily such a very interesting and comprehensive hypothesis as that referred to.

Mr. Hehner said that it was often difficult to distinguish between the alkalinity produced by sodium carbonate or calcium carbonate and the alkalinity produced by magnesium salts. The waters now under consideration, however, contained rather small quantities of magnesium and calcium, so that the alkalinity would not be much affected by these, but would be due practically to sodium carbonate. The question of the presence of sodium carbonate in chalk was a very interesting one. Chalk was a marine deposit, and the sea certainly contained no sodium carbonate, because it contained large quantities of magnesium chloride, and the two were incompatible. Consequently, the chalk originally was not an alkaline deposit in the sense of containing sodium carbonate, and the question as to how it became alkaline was a difficult one to answer. Anybody who had looked at a chalk cliff on the South Coast must have noticed the manner in which the alternating deposits of chalk and flint occurred therein. He had discussed the subject with many geologists,

but had never been able to obtain a feasible explanation of the fact that there occurred in the chalk at every 15 or 18 inches a well-defined layer of flint, then a layer of chalk without any flint, and so on. He did not think it would satisfy anyone to be told that during a certain era there was a deposit of chalk, that then the circumstances became such that foraminifera became less abundant and sponges predominated, these again dying out and giving place to a deposit of chalk. He had always thought that, in addition to the action of living organisms, chemical action must have played a considerable part in producing the alternation of deposits, and it was quite obvious from Mr. Fisher's observations that a chemical alteration had taken place in the chalk. Chalk contained appreciable quantities of silicates, and silicates were readily dissolved, but silica was, on the other hand, readily deposited from water. Rain-water percolating through the soil becomes saturated with silica, which readily deposits again round nuclei (spicules of sponge) and flint forms. Supposing that the chalk also contained sodium salts, then it was probable that, with water containing carbonic acid, decomposition of the silicates would take place with the formation of sodium silicate, and that through the action of this on the calcium carbonate sodium carbonate might be formed. He had always attributed the alkalinity of the water from the chalk below the London clay to the Thanet sand, or possibly to the greensand. But the explanation suggested by Mr. Fisher would be far more satisfactory if it were confirmed by an examination of the waters from other chalk districts where similar conditions existed. No doubt in this or other countries there were many beds similarly situated.

Mr. Fairley said that in certain parts of Yorkshire alkaline waters occurred very frequently, especially in the neighbourhood of coal formations, and under certain portions of the millstone grit and other formations. He was not sure whether such waters always occurred in connection with beds of clay, but the quantity of carbonate of soda which they contained was sometimes very considerable. The question had sometimes occurred to him whether this sodium carbonate might have been formed from sodium salts in some way similar or analogous to that in which sodium nitrate was formed in the nitrate beds of Chili or sodium carbonate in the natron beds of Egypt, where organic matter played some part in the changing of what was no doubt originally sodium chloride into sodium nitrate in some cases, and into sodium carbonate in others. He thought that in the case of waters like those referred to it would be interesting to look for other alkaline bases. In the colliery waters of the North of England lithia was nearly always present, though it had been reported in only a few instances. Its presence was probably attributable to the decomposition of lithia mica occurring in the coal measures. In examining water for lithia, it was impossible to obtain satisfactory results without subjecting the water to a very careful process of concentration as a preliminary to spectroscopic examination. The leading principle should be to use only such reagents, and a minimum quantity of them, as would introduce the least possible risk in the analysis by introducing the substance that was being sought for. In the case of a water containing lithium, just sufficient barium hydrate would be added to precipitate all the alkaline earth metals, with as little excess as possible; and the residue obtained, after elimination of the excess of barium, treated with a mixture of absolute alcohol and ether. On evaporation a

residue was obtained, frequently deliquescent after exposure, which showed the characteristic lithium bands very distinctly, when such bands could not be seen by direct spectroscopic examination of the solid residue.

Mr. Blount said that he could quite confirm Mr. Fairley's remarks in reference to lithia. Lithia was a very common constituent of natural water, in almost any sample of which he would undertake to find it by the method described by Mr. Fairley.

Mr. Chapman said that in the determination of alkaline carbonates by evaporating to dryness and extracting with water it was necessary, in the case of waters containing appreciable quantities of magnesium salts, to take care that the washing was not carried too far, or the results obtained would be too high, owing to the distinct solubility of magnesium carbonate. With regard to the suggestion that carbonate of soda could not coexist in water with lime salts, if bicarbonate of soda solution were added to water containing several grains per gallon of sulphate of lime or sulphate of magnesia, the water remained perfectly bright. The conclusion he had come to was that those who systematically made complete mineral analyses of water observed a good many facts not commonly stated in the text-books. He preferred to express the results of such analyses, not in terms of the water, but in terms of the residue. He would like to ask Mr. Fisher whether he had formed any hypothesis to explain the presence of what was an extremely common constituent of pure waters of the class referred to, namely, saline ammonia in quantities, which in any other waters would be held to indicate pretty extensive pollution. He was not quite at one with Mr. Fisher in thinking that these waters rarely contained any appreciable quantity of lime or magnesia. In his experience, waters from similar beds in other parts of the country, while they contained often not more than 2 or 3 grains per gallon of the carbonates of lime and magnesia, sometimes contained as much as 8 or 10 grains per gallon (but not often more) of calcium carbonate.

Mr. C. A. Mitchell inquired at what temperature the total solids were dried. It was rather difficult to expel all moisture at the ordinary temperature of the water-bath, while at a higher temperature magnesium chloride would be decomposed.

Mr. Fisher said that the total solids were dried on the water-bath.

The President (Dr. Voelcker) said that, to sum up Mr. Fisher's conclusions, they were to the effect that, in the case of the bare chalk, the alkaline carbonates were practically washed out entirely; but in the case of the chalk being covered by clay there was not such a thorough washing out, carbonate of soda being present to a very much larger extent than carbonate of lime, the carbonate of soda being originally an integral constituent of the chalk. This matter, of course, had to be looked at from several points of view and in regard to many different kinds of water. Analysts had been in the habit of saying that waters drawn from the chalk, but which were found to contain only a little lime, were to be regarded as somewhat abnormal, the explanation usually suggested being that there was some infiltration from the sea, and that chemical changes took place which resulted in the formation of carbonate of soda. He had before him some analyses of waters which would very well compare with the water which Mr. Fisher had referred to, namely, those of two supplies from deep artesian wells, situated respectively at the Royal Hotel, near Blackfriars Bridge, and

in Stamford Street, near Waterloo Station. These waters, which he had analysed at different times, contained 50 grains of total solids per gallon, only 1 or 2 grains of lime, about ½ grain of magnesia, and 18 to 20 grains of carbonate of soda, together with large amounts of sulphate of soda; these results would lead one to similar conclusions to those which Mr. Fisher had arrived at. But there were other cases which did not quite bear this out, and he agreed with Mr. Chapman that there was not sufficient evidence of the non-existence of alkaline carbonates and lime salts together. In the case of a water from Newington, London, containing 86 grains of total solids per gallon, there were over 15 grains of lime per gallon, with large quantities of alkaline carbonates and alkaline sulphates as well. In some districts of Yorkshire, near Malton, for example, there were limited areas where deposits of carbonate of soda occurred, and the waters from the chalk in those neighbourhoods, although there was no overlying clay, frequently contained as much as 150 grains per gallon of carbonate of soda. He had made inquiries in the neighbourhood, and of geological authorities, but had never been able to find any exact records of the occurrence of these deposits or any satisfactory explanation of their presence.

Mr. Dibdin thought that probably biology would afford some clue to the means by which the carbonate of soda was produced. When the chalk was deposited it originally contained a certain quantity of organic matter derived from marine organisms, and it seemed quite possible that, as the result of the microbic action that would take place during the infiltration of an aerated water through the chalk, the carbon of this organic matter had become converted into carbonic acid, and under these bacterial conditions it might have been that some of the original sulphate of soda underwent change, resulting in the formation of sulphate of lime and carbonate of soda.

Mr. Fisher said that he had been thinking over this matter for years, and had rejected various explanations, until he obtained the direct evidence of the presence of carbonate of soda in the chalk, which seemed to decide the question. He had come to the conclusion that such was the probable explanation before he was able to obtain practical evidence of it. He hoped to be able to get evidence that the limestones in the other beds also contained free alkaline carbonates, which he thought would add to the strength of the position. What in the tables he had termed "alkalinity" represented all the carbonates as carbonate of soda. The figures give some information as to the relative alkalinity of the samples. A hundred c.c. of ordinary chalk-water generally neutralized about 5 c.c. of decinormal acid, but some of these samples neutralized 7, 8, 9 or 10 c.c., which afforded an idea of the amounts of alkaline material present, translated from decinormal acid to carbonate of soda. He was quite familiar with the district referred to by Mr. Richmond. The gault clay which occurred in the neighbourhood, and which varied in width from two to four miles, was a great difficulty in the way of getting proper water. A supply by gravitation was difficult owing to the distance from the outcrop of the chalk and the greensand, while if the clay were pierced the water obtained was extremely saline. The water from a deep well at Swindon was so saline that as it came down the Thames, at times, it affected the chlorine in the Oxford water-supply. This saline water had been traced by him in the Thames up to Lechlade, and from thence to Swindon by

Mr. Groves. On the 1 inch to the mile geological Ordnance map it would be seen that there was good reason for the variations in the quality of the water coming from below the gault. In some places the collecting-ground was large, and consisted to a considerable extent of exposed beds of coral rag, and the water came down gradually beneath the hills. In other places the Oxford clay covered the porous collecting-ground, and there was only a very narrow strip of pervious rock which could get any water at all. Shrivenham Station was just on the edge of the gault, and a little further north the coral rag was exposed, and there no excess of saline matter was to be observed. If the explanation were correct as to the cause of the alkalinity, it ought to apply to any similar bed of chalk below the London clay. There was such a bed of London clay in the southern part of Hampshire, near the New Forest, on both sides of which the chalk downs rose to a considerable height, and he hoped to have an opportunity of examining a sample from this locality. All these saline waters contained exceptional quantities of ammonia, and some contained considerable quantities of organic matter. At Marsh Gibbon, where the water flowed out of a borehole, it came out quite brown, from the quantity of organic matter dissolved in passing under the ground for three miles; that water contained also a notable quantity of free ammonia. Where ammonia was present in these waters there was very seldom any quantity of nitrates. The question of the solubility of calcium carbonate in waters containing an excess of free ammonia was dependent to some extent upon the proportion of carbonic acid present. In a soda-water made from a hard water considerable quantities of carbonate of lime would be found in solution as well as carbonate of soda. It was generally the case that those waters which were alkaline contained very small proportions of carbonate of lime, though he was not certain as to how the lime might have disappeared. Frequently there was only so much lime in solution as would dissolve *per se* as calcium carbonate.

ABSTRACTS OF PAPERS PUBLISHED IN OTHER JOURNALS.

FOODS AND DRUGS ANALYSIS.

Methods for the Detection and Estimation of Arsenic in Beer, Brewing Materials, and Food. W. Thomson and J. P. Shenton. (*Journ. Soc. Chem. Ind.*, 1901, xx., 204.)

Marsh-Berzelius Test.—Fifty c.c. of *beer* are placed in an 8-ounce flask, and heated to the boil with agitation, then evaporated to a syrup on a sand-bath; 20 c.c. of nitric acid and 5 c.c. of strong sulphuric acid are next added and the liquid boiled down till it darkens. One or 2 c.c. more of nitric acid are introduced as often as discoloration occurs until the organic matter is completely destroyed (about 25 c.c. are usually required). The liquid finally occupies a volume of about 5 c.c.; it is diluted with an equal volume of water and boiled, to expel nitrous and nitric acids.

It is then made up to 15 c.c. The Marsh apparatus consists of a 50 c.c. conical Jena flask fitted with a wooden stopper, made tight with paraffin wax. It is conveniently provided with a graduated stoppered tube-funnel, holding 15 c.c., to contain the arsenical liquid, the stop-cock of which is so arranged as either to permit introduction of measured quantities of the arsenical liquid, or of dilute sulphuric acid (1 : 7). The flask is filled one-third full of granulated zinc. The hydrogen passes through a tube containing a coil of filter-paper and a plug of cotton-wool, then enters a hard glass tube $\frac{3}{16}$ inch in diameter, drawn out at the end. The acid should be added so that the hydrogen flame is between $\frac{3}{32}$ and $\frac{1}{8}$ inch long; and the Bunsen burner is placed where it will heat $\frac{3}{4}$ inch of the tube to redness about $\frac{3}{16}$ inch from the constriction. Sufficient of the arsenical liquid is gradually introduced (1, 2, or 3 c.c.) to yield a mirror of standard size, and thus the process becomes quantitative. Standard mirrors are prepared by adding from weak standard arsenious acid solutions quantities corresponding to the $\frac{1}{300}$, $\frac{1}{400}$, $\frac{1}{500}$, $\frac{1}{700}$, $\frac{1}{1000}$, $\frac{1}{750}$, $\frac{1}{600}$, and $\frac{1}{850}$ grain per gallon. In making the comparisons, the flask should be as small as possible, and all dimensions and weights, etc., should be kept identical. Pure sulphuric and hydrochloric acids are obtained by distillation with a little chromic acid, chlorine in the HCl being destroyed by pure phenol. Rod zinc should not be used, and the metal should be quite pure, since various impurities, not all of which are known, tend to hold back the arsenic. Additions of platinum, platinic chloride, aluminium, magnesium, and lead or zinc salts, should be avoided.

In the case of *malt*, 5 grammes are treated with 25 c.c. of nitric acid, heated cautiously, and 5 c.c. of sulphuric acid added, proceeding then as above. With *coal* and *coke*, 5 grammes of the former, or 2·5 grammes of the latter, in fine powder, are heated with 25 c.c. of nitric acid for one hour, then with 5 c.c. of sulphuric acid for half an hour. After cooling, the mixture is diluted to about 100 c.c. and filtered. The filtrate is concentrated with more nitric acid to destroy organic matter, and finally evaporated free from nitrous and nitric acid. The resulting solution is then examined for arsenic as above.

Reinsch Test.—The present authors cannot accept the process specified by the Brewers' Association (this volume, p. 13): (1) Because it is less delicate; (2) because it is less distinct than the Marsh test, the colour of the sublimate obtained in the Reinsch test not being so convenient for comparison as the Marsh black; (3) because the whole of the arsenic is not recovered from the copper by heating; (4) because some arsenic will be volatilized when the beer is concentrated with hydrochloric acid; (5) because arsenic acid may be present and delay the reaction; and (6) because the small piece of copper is not likely to collect all the arsenic present in the liquid.

Marsh-Gutzeit Test.—This, as modified by Tyrer, was tested. The arsine was passed through potash bulbs charged with strong lead acetate solution before being applied to the paper treated with mercuric chloride. The gas does not appear to dissolve in the solution to any appreciable extent; and the authors consider this the best method of working the test, questioning whether the Marsh-Gutzeit is not the best for the approximate determination of minute quantities of arsenic after the organic matter of the original sample has been destroyed. One disadvantage is that the yellow colour cannot be seen by gaslight.

Neither sulphuric nor hydrochloric acid free from chlorine removes arsenic from arsenical glass when stored therein for seventeen hours in the cold. Sodium selenite yields a dark-brown precipitate in the Marsh flask, and gives a bright red mirror. No indication of that element was discovered in any of the beers and sugars examined. F. H. L.

ORGANIC ANALYSIS.

A Gas-volumetric Method of Determining Formaldehyde. E. Riegler. (*Zeit. anal. Chem.*, 1901, xl., 92-94.)—If hydrazine sulphate be treated with a solution of iodic acid, the whole of its nitrogen is liberated, as in the equation :

$$5(N_2H_4H_2SO_4) + 4HIO_3 = 5N_2 + 12H_2O + 5H_2SO_4 + 4I.$$

If, however, formaldehyde be present, it combines with the hydrazine to form a hydrazone, which is only decomposed after some time by the iodic acid. Thus, a quantity of hydrazine corresponding to the hydrazone remains undecomposed, and the volume of nitrogen is smaller.

The reagents required are a solution of 1 gramme of hydrazine sulphate in 100 c.c. of water, and a solution of 5 grammes of crystallized iodic acid in 50 c.c. of water.

The apparatus used by the author was the well-known Knop-Wagner nitrometer. Twenty c.c. of the hydrazine solution are mixed with 20 c.c. of water in the evolution flask, and 5 c.c. of the iodic acid solution introduced into the inner vessel. The nitrogen is collected and measured in the usual manner.

A definite quantity of the aqueous solution of formaldehyde containing not more than 0·08 gramme of that substance is now shaken with 20 c.c. of the hydrazine sulphate solution, and the mixture left for fifteen minutes. It is then introduced into the evolution flask of the apparatus, together with about 10 c.c. of water, and the nitrogen liberated as before by the introduction of 5 c.c. of the iodic acid solution, the shaking being continued for half a minute, and the reading taken after two minutes (not more).

The difference between the two results, calculated to standard temperature and pressure, and multiplied by the factor 2·7, gives the corresponding amount of formaldehyde in milligrammes. C. A. M.

The Separation and Determination of Ammonia and Methylamines. H. Quantin. (*Ann. de Chim. anal.*, 1901, vi., 125, 126.)—This method is based on the following principles : If a solution containing both free or combined ammonia and methylamines be treated with sodium phosphate and then with magnesium sulphate, the whole of the ammonia is precipitated as ammonium magnesium phosphate, provided the liquid be alkaline from the presence of methylamines, which should be added if necessary.

An aliquot part of the solution is diluted to a litre :

(*a*) The total volatile bases are determined in 100 c.c. of this solution, and calculated as ammonia = A.

(*b*) One hundred c.c. of the solution are treated with sufficient sodium phosphate and magnesium sulphate to precipitate a quantity of ammonia equal to A. Purified methylamines are added to insure the complete precipitation, and the precipitate allowed to settle for six hours, after which it is collected and washed with water until the washings are neutral.

(*c*) The precipitate and filter-paper are transferred to a distillation flask, the precipitate dissolved in hydrochloric acid, an excess of sodium hydroxide added, and the ammonia distilled into standard acid in the usual manner.

(*d*) One hundred c.c. of the original solution are concentrated to a small volume, the total ammonia and mono- and di-methylamines precipitated with platinum chloride, and the precipitate collected, dried, and weighed. Let p represent the weight of the precipitate, and p_1 the weight of platinum obtained on igniting the precipitate. Seventeen parts of ammonia, 31 parts of monomethylamine, and 45 of dimethylamine yield respectively 223, 237, and 251 parts of platinochlorides, containing 98·5 parts of platinum.

If the quantities of the three bases be represented by x, y, z, we obtain the equations:

$$(1)\quad \frac{223}{17}x + \frac{237}{31}y + \frac{251}{45}z = p\,;$$

and

$$(2)\quad \frac{x}{17} + \frac{y}{31} + \frac{z}{45} = \frac{p_1}{98{\cdot}5}.$$

Substituting for x the result from the determination of ammonia, two equations are obtained, giving the values of y and z.

Finally, the difference $A - (x + y + z) = P$ gives in terms of ammonia the proportional amount of trimethylamine, of which the real quantity will be $\frac{P \times 59}{17}$.

[Ammonium magnesium phosphate does not readily yield its ammonia on mere distillation with sodium hydroxide.—B. D.]

C. A. M.

Quantitative Determination of Malic Acid. A. Hilger. (*Zeit. für Untersuch. der Nahr. und Genussmittel*, 1901, iv., 49, 50.)—The author proposes a method based upon the observed fact that 1 gramme of malic acid, on boiling with a slightly alkaline or neutral solution of palladium chloride reduces 0·294 gramme of palladium. One hundred c.c. of the liquid—*e.g.*, wine—under examination are evaporated, to remove volatile constituents, and the residual liquid made feebly alkaline with basic lead acetate. The resulting precipitate, which contains the whole of the malic acid, is filtered off, washed with cold water, and dissolved in a little dilute boiling acetic or nitric acid. Sodium carbonate is then added to the boiling solution to alkaline reaction, and a current of carbonic acid gas passed through the liquid. The precipitate of basic lead carbonate is removed by filtration, the filtrate concentrated to 100 c.c. It is then neutralized with hydrochloric acid, transferred to a 500 c.c. Erlenmeyer flask, treated with 10 c.c. of a 5 per cent. palladium chloride solution, and boiled for ten minutes. The reduction of the palladium chloride is accompanied by a brisk disengagement of carbonic acid gas; when this ceases, the liquid is made faintly acid with hydrochloric

acid, and the heating continued until the palladium has consolidated and settled to the bottom of the vessel. The precipitated metal is then filtered off, washed, dried, heated in a stream of carbonic acid gas, cooled, and weighed. H. H. B. S.

Remarks upon the Halphen Reaction with Cottonseed Oil. E. Wrampelmeyer. (*Zeit. für Untersuch. der Nahr und Genussmittel*, 1901, iv., 25.)—The author considers that Soltsien's modification of Halphen's procedure (*Zeit. öffent. Chem.*, 1899, 106, 135) impairs the sensitiveness of the test. According to his experience, a decided indication can be obtained in a quarter of an hour by the use of amyl alcohol, as stated by Halphen, whereas by Soltsien's procedure three-quarters of an hour are required. He, however, does not consider it necessary to revert to the use of the salt-water bath. The author's method of working is as follows: A moderately thick test-tube, about 2·5 centimetres wide and 15 centimetres long, and graduated at 10 centimetres, is filled to the graduation with the oil to be examined, and an equal volume of amyl alcohol and about 2 c.c. of a 1 per cent. solution of sulphur in carbon disulphide added. A bored cork with fitted tube is then inserted in the neck of the test-tube, and the whole heated in a boiling-water bath for a quarter of an hour (*cf.* ANALYST, 1898, 131; 1899, 214; 1900, 105). H. H. B. S.

The Optical Examination of Fats and Waxes. G. Marpmann. (*Chem. Rev. Fett. u. Harz. Ind.*, 1901, viii., 65-68.)—Although the refractive index of beeswax affords a rapid means of judging as to its impurity, there has hitherto been the drawback that the refractometer is injured by being brought to the necessary temperature of 70°-80° C.

This can be obviated, however, by making use of the fact that the refractive index of a mixture is equal to the arithmetical mean of the indices of its components. Thus, by mixing beeswax with a fatty oil, a mass which remains fluid at 40° C. can be obtained, and its refraction index readily determined.

Ethereal oils are also more or less suitable substances for the admixture, but ether, chloroform, or higher alcohols, benzene, etc., are less satisfactory. The author gives the results of experimental tests with mixtures of beeswax with peppermint oil and clove oil. C. A. M.

Determination of Cresol. F. Russig and G. Fortmann. (*Zeits. angew. Chem.*, 1901, 157.)—The following method gives somewhat higher results than that of Raschig (ANALYST, 1900, 298). 50 grammes of the cresol are weighed into an Erlenmeyer flask, and 125 grammes of sulphuric acid of 66° B. are added. The mixture is warmed to 60° or 70° C., and allowed to stand one or two hours. The nitrating apparatus consists of a glass retort of about 1 litre capacity fitted with a dropping funnel. Into the retort 400 grammes of nitric acid of 40° B. are introduced and heated to 60° C., the flame being then removed. The sulphuric acid is allowed to drop into the nitric acid. This takes one and a half to two hours. About twenty minutes after all has run in, the contents of the retort are poured into a large porcelain basin containing 200 c.c. of water, and a further 200 c.c. of water are used

to wash out the retort. It is then allowed to stand overnight. The trinitro-cresol is then collected on a hardened filter-paper, washed with 200 c.c. more water, dried at 95° to 100°, and weighed. Pure metacresol thus yields 175·6 per cent. of trinitro-cresol as against 174 per cent. given by Raschig's method.

The authors find that the method of Ditz and Cedivoda (*Zeits. angew. Chem.*, 1900, 1050) is quite unreliable. A. M.

Determination of Cresol. H. Ditz. (*Zeits. angew. Chem.*, 1901, 160.)—The author replies to the criticisms of Russig and Fortmann upon his method (see above). A. M.

INORGANIC ANALYSIS.

The Estimation of Manganese in Ferro-Chromium Alloys. J. T. (*Chem. News*, vol. lxxxiii., p. 25.)—One gramme of the powdered alloy is fused with several grammes of sodium peroxide in a nickel crucible, then dissolved in water, boiled for five to ten minutes, and filtered. The mixed oxides of iron, manganese and nickel filter badly, but settle readily, and should therefore be boiled repeatedly with very dilute ammonium carbonate, and collected on the filter when the liquid has nearly or quite lost its yellow colour. The residue and (asbestos) filter are treated with equal parts of nitric acid (1·20) and water, and sufficient ferrous sulphate to fully reduce the nickel and manganese oxides. The solution is thoroughly cooled, oxidized with sodium bismuthate, and the filtered permanganate titrated with hydrogen peroxide, according to the instructions of Reddrop and Ramage.

The green colour of the titrated solution, due to nickel from the crucible, is destroyed by a slight excess of permanganate. C. S.

On Vogel's Method for the Qualitative Detection of Cobalt. F. P. Treadwell. (*Zeit. anorg. Chem.*, xxvi., 108.)—This method is especially suitable for the detection of traces of cobalt in nickel salts, even in the presence of iron. To the solution of the nickel salt a concentrated solution of ammonium thiocyanate is added, and the whole shaken with a few c.c. of amyl alcohol. If the upper alcoholic layer remains colourless, neither iron nor cobalt is present; if it becomes red, iron is indicated. In this case a few drops of sodium hydroxide solution are added, and the liquid is again shaken; if cobalt is present, the upper layer will now be coloured blue, through the formation of a compound $Co(NH_4)_2(CNS)_4$. A. G. L.

Estimation of Aluminium in Steel. E. Spatz. (*Zeits. öffentl. Chem.*, 1901, vii., 60; through *Chem. Zeit. Rep.*, 1901, 79.)—One gramme of filings is dissolved in 2 c.c. of strong sulphuric acid and 15 or 20 c.c. of water. The solution is transferred to a silver basin, treated with 40 grammes of crystallized ammonium oxalate, and electrolyzed. [No electrical measurements are given.] At intervals a few drops of the liquid are withdrawn through a fine pipette, evaporated and ignited on platinum foil, and tested for iron. When only traces thereof remain, the main

solution is run into another vessel, and the former basin is rinsed out and rubbed well with a rubber-tipped glass rod. The liquid is evaporated to dryness, brought to a dull red heat, and fused with potassium bisulphate. The melt is dissolved in water, acidulated with sulphuric acid, mixed with some ammonium phosphate, and precipitated with ammonia; the precipitate is collected, dissolved in hydrochloric acid, thrown down again with ammonia, filtered, dried, and weighed as iron plus aluminium phosphate. It is then once more fused with bisulphate, the melt taken up in dilute sulphuric acid, a weighed piece of pure zinc added, and the iron titrated with permanganate. By calculating the weight of iron found into phosphate, and deducting from the mixed phosphates, the amount of aluminium phosphate derived from the steel is left. If the sample only contains traces of aluminium, 4 or 5 single grammes should be electrolyzed separately, and the united liquids examined as above.

F. H. L.

Calcium Compounds in Soil. Estimation of "Available" Lime. D. Meyer. (*Landw. Jahrb.*, 1900, xxix., 913; through *Chem. Zeit. Rep.*, 1900, 377.)—The author describes an investigation of twenty-six different specimens of soil, on twenty-one of which experiments with growing plants were also carried out. The total lime in the samples varied from 0·092 to 1·271 per cent. The average proportion of lime in the light soils was 0·333 per cent., in the heavy soils 0·694 per cent.; but although the latter usually contained more than the former, this was not always the case. In twenty-two of the samples the proportion of carbon dioxide ranged from 0·020 to 0·076 (mean 0·045) per cent.; four samples contained from 0·168 to 0·350 per cent. of CO_2. The average for the light soils was 0·052 per cent., for the heavy soils 0·098 per cent. Most of the soils contained less than 25 per cent. of their total lime as carbonate; there was a larger proportion of carbonates in the light than in the heavy soils. Samples having the same proportion of carbon dioxide contained totally different proportions of lime.

The solubility of the lime in dilute hydrochloric acid varied considerably; in light soils it was less soluble than in the heavy soils. There was no connection between the solubility of the lime in dilute hydrochloric acid and the proportion of very fine particles in the soil. There was no connection between the amount of carbon dioxide, of sulphuric acid, or of phosphoric acid and the solubility of the lime. Magnesia stood in much the same relation as the lime. Judging from the small proportion of carbon dioxide, sulphuric acid, and phosphoric acid, it is clear that most of the lime in these soils existed as silicates. Since the solubility of the lime in dilute acid was less in the light than in the heavy soils, the latter must contain their lime chiefly as easily decomposable silicates.

Different compounds of lime exhibit notable differences in their effects upon the growth of plants. The useful compounds of calcium are, according to the author, the carbonate, the sulphate, and the easily decomposable silicates; and as they can be dissolved by a neutral solution of ammonium chloride or nitrate, by adopting these solvents it becomes possible to devise a process for estimating available lime which gives figures agreeing better with the results of experiments on plants than when hydrochloric acid is employed.

Meyer accordingly proposes to estimate the "available" lime by digesting the sample of soil for three hours on the water-bath with a 10 per cent. solution of ammonium chloride, determining the calcium in the filtrate without removing the silica. A yield of 0·25 per cent. by this process may be considered normal; it should not fall below 0·20 per cent. F. H. L.

On the Quantitative Determination of Magnesium by Means of Organic Bases. W. Herz and K. Drucker. (*Zeit. anorg. Chem.*, xxvi., 347.)—Magnesium can be separated from the alkali metals in the same way as zinc, by precipitating the solution in the cold with an aqueous solution of dimethylamine; after standing for several hours the precipitated magnesium hydroxide is filtered off, washed rapidly with water containing dimethylamine, and ignited over a Teclu burner. The results obtained were satisfactory. The separation may also be effected by means of guanidine, obtained from the carbonate by adding an amount of barium hydroxide insufficient to liberate all the guanidine. The carbonate itself does not precipitate magnesium completely. A. G. L.

Determination of Fluorine in Zinc Blendes. F. Bullnheimer. (*Zeits. angew. Chem.*, 1901, 101.)—The apparatus used consists of a flask of about 300 c.c. capacity fitted with a rubber stopper, through which pass inlet and outlet tubes and a thermometer. The outlet tube leads first through a U-tube filled with glass wool, then through a cooled Winkler coil into a Drehschmidt wash bottle containing potassium chloride solution. About 2·5 grammes of the blende are weighed out, mixed with 3 to 5 grammes of quartz powder, and introduced into the flask. A mixture of 20 grammes of chromic acid and 100 c.c. of concentrated sulphuric acid are added, the flask is closed, shaken round, and purified air is led through the apparatus. After a time the temperature of the contents is gradually raised to 80° C. An energetic reaction then sets in, and the flame is removed. When this has ceased, the temperature is raised to 150°, and the heating is continued for three hours. The whole of the fluorine will then have been driven off, and will have liberated acid according to the equations:

$$3SiF_4 + 2H_2O = 2H_2SiF_6 + SiO_2$$
$$2H_2SiF_6 + 4KCl = 2K_2SiF_6 + 4HCl.$$

The acid is titrated with decinormal alkali, using phenolphthalein as indicator. A. M.

A New Test for Chlorine for Use with the Blow-pipe. Henry W. Nichols. (*Amer. Chem. Journ.*, xxv., 315.)—The substance to be examined is heated with potassium bisulphate in a tube closed at one end and a piece of filter-paper moistened with a solution of cobalt nitrate held at the mouth of the tube. If a chloride is present, the paper will turn blue; if a bromide or iodide, green. In the case of small quantities of chlorides, it may be necessary to dry and warm the paper to produce the colour. If both chlorides and bromides are present, the paper turns blue first, and then green, or, if but little bromine is present, bluish-green. A mineral con-

taining about 2·2 per cent. of chlorine gave a good reaction by this method, but the test is much less sensitive for bromides than for chlorides.

Chlorides in solution may be detected by dipping a piece of filter-paper into the liquid, adding a drop of cobalt nitrate, and drying the paper, but the test is not very delicate.

The chlorochromic acid test may be modified by carrying it out in a plain test-tube, and holding a piece of filter-paper moistened with a saturated solution of sodium carbonate at the mouth of the tube. The paper retains the chromium, which can be easily identified in the ash. A. G. L.

Proximate Analysis of Clays. W. Jackson and E. M. Rich. (*Journ. Soc. Chem. Ind.*, 1900, xii., 1087.)—In order to estimate the amount of clay substance, felspar, and quartz in a sample of clay, it is usual to utilize the action of sulphuric acid on hydrated silicate of alumina (clay), which brings about the solution of the alumina as sulphate. The silica remains in the amorphous condition, and is subsequently taken up with dilute alkali. The authors find, however, that both acid and alkali act upon felspar, dissolving it as a whole, especially in the case of plastic clays. A. M.

On the Persulphates and their Application in the Laboratory and Industries. R. Namias. (*L'Orosi*, xxiii., 218, through *Zeit. anorg. Chem.* xxvi., 125.)—For the analysis of persulphates, the salt should be allowed to stand in the cold for twelve hours with potassium iodide solution, and the liberated iodine titrated. As solutions of persulphates remain unchanged for a long time, they can be used in volumetric analysis. Ammonium persulphate precipitates manganese dioxide quantitatively from hot ammoniacal solutions of manganese salts. When used in excess in hot solutions, it is also capable of oxidizing chromic oxide to chromic acid in one hour. A. G. L.

The Recalculation of the Chemical Analyses of Rocks. J. F. Kemp. (*School of Mines Quart.*, xxii., 75.)—If the chemical analysis of a rock is accompanied by a microscopical study of thin sections of the rock, it is often possible to calculate the percentages of the several minerals present. In the first place, the "molecular ratios" of the various molecules present are determined by dividing the percentage values found by the corresponding molecular weights. The following example of a syenite is taken from a report by W. H. Weed and L. V. Pirsson (Twentieth Annual Report Directory United States Geological Survey, iii., 466), the molecular proportions being given under the respective percentage values:

SiO_2.	Al_2O_3.	Fe_2O_3.	FeO.	MgO.	CaO.	Na_2O.	K_2O.	P_2O_5.	Cl.	Total.
64·64	16·27	2·42	1·58	1·27	2·65	4·39	4·98	0·37	0·05	98·62
1·077	0·158	0·015	0·022	0·031	0·047	0·070	0·053	0·002	0·0014	

The remaining 1·50 per cent. consisted of TiO_2, H_2O, CO_2, BaO, SrO, and is neglected in the recalculation. An examination of thin sections with the microscope showed the minerals present in the rock to be the following: Orthoclase, K_2O, Al_2O_3,

$6SiO_2$; plagioclase, albite Na_2O, Al_2O_3, $6SiO_2$, and anorthite CaO, Al_2O_3, $2SiO_2$; hornblende, MgO, SiO_2, CaO, SiO_2 and FeO, SiO_2; magnetite, Fe_2O_3, FeO; apatite, $9CaO$, $CaCl_2$, $3P_2O_5$; quartz, SiO_2.

On inspecting the above formulæ, it is evident that all the K_2O (53) is in the orthoclase, all the Na_2O (70) in the albite, and the remaining Al_2O_3 (35) in the anorthite, requiring an equal number of molecules of CaO (35). The rest of the CaO (47) is in the hornblende and apatite, and the apatite can be calculated from the P_2O_5 (2). All the Cl_2 (·7) is in the apatite, and the MgO (31) in the hornblende. All the Fe_2O_3 (15), and an equivalent amount of FeO (15), are in the magnetite, the remaining FeO (7) being in the hornblende. The excess of SiO_2 constitutes the quartz.

To obtain the percentages of the minerals in the rock, the molecular proportion of one of the constituents of the mineral is multiplied by the molecular weight of the mineral—*e.g.*, the percentage of orthoclase is $0{\cdot}053 \times 556 = 29{\cdot}46$. Thus the following percentages are obtained: Magnetite 3·50, hornblende 4·60, anorthite 9·73, albite 36·68, orthoclase 29·46, quartz 13·54, apatite 1·10: Total, 98·61.

The ratio of the albite molecules to those of anorthite is as 4 to 1. From the optical properties of the plagioclase, Pirsson found it to be approximately Ab_2An. Consequently half the albite is in the orthoclase, and the ratio of alkali felspar to soda-lime felspar is as 424 Or + 280 Ab to 280 Ab + 140 An, or nearly as 5 to 3. By means of a result of this nature, a rock may be readily classified as an orthoclase or plagioclase rock.

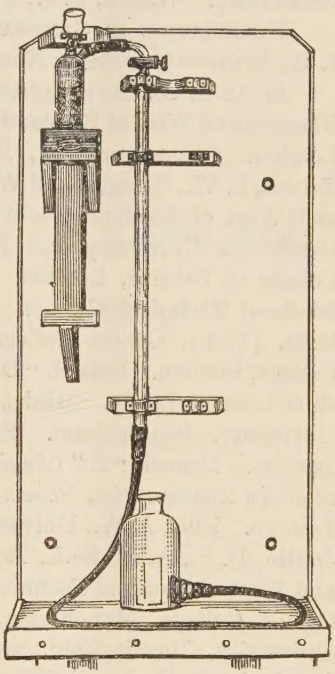
In some cases, however, the recalculation is not so simple as in the above example, and it may become necessary to separate and analyse one of the minerals before proceeding with it. The author gives the formulæ of the more important minerals, and also a set of tables giving the molecular ratios corresponding to different percentages of various oxides. A. G. L.

APPARATUS.

A Flask Support. F. Reiss. (*Chem. Zeit.*, 1901, xxv., 351.)—This device consists of a round-topped tripod carried on a short sleeve, which fits over the tube of a Bunsen burner, and which bears a thumbscrew for clamping it at the desired height. The upper ring also supports a suitably bent wire to hold the neck of a flask whilst being heated in a sloping position, as is necessary in working the Kjeldahl process (for which the stand is primarily intended). Obviously the neck support can be omitted and the sleeve increased in length indefinitely; and thus the apparatus forms the only stand required to hold a flask or retort before the mouth of a condenser, the gas flame burning at the upper orifice of the sleeve itself.

F. H. L.

A Form of Volumenometer. F. W. Jones. (*Chem. News*, vol. lxxxiii., p. 100.)—The apparatus is designed for specific gravity determinations in the case of substances which cannot be brought into contact with water, and consists of a gas burette with stopcock, connected at one end with a stoppered bottle, and at the other end with a reservoir of mercury. The bottle should be small, and have a stopcock fitted to one of the two openings in the ground stopper. It is mounted on a spring platform, whilst the stopper is fixed, and the second opening is connected, by a rubber tube, with the burette, which, in turn, is connected to the mercury reservoir as usual.



A known weight of substance being placed in the bottle and the stopcocks opened, the reservoir is raised until the mercury just reaches the top of the burette; whereupon the burette stopcock is closed and the bottle placed on the lower shelf, the barometric pressure of the mercury being then read off. When the air in the bottle has attained room temperature, the bottle stopcock is closed and that of the burette opened, the fall of the mercury in the burette and its rise in the reservoir being noted.

The results are calculated by the formula: $V_1 = v\frac{d+r}{P-d-r}$ wherein V_1 is the gas volume of the bottle and connections; P = the pressure; v the volume of gas in the burette at the reduced pressure; d the amount the mercury falls in the burette; and r the amount of its rise in the reservoir. Then, V being the volume of bottle and connections when empty, the specific gravity is $\frac{W}{V - V_1}$. C. S.

INSTITUTE OF CHEMISTRY OF GREAT BRITAIN AND IRELAND.

THE following is a list of the names of the candidates who passed the examination of the Institute of Chemistry held in July, 1901:

INTERMEDIATE EXAMINATION.—Armstrong, A. G., Glasgow and West of Scotland Technical College, and at King's College, London. Arnaud, F. W. F., King's College, London, also under Messrs. W. F. Lowe and C. H. Cribb, FF.I.C. Brooker, J. W. G., Finsbury Technical College, London. Finnemore, H., Pharmaceutical Society's Laboratory, and at King's College, London, Green, G. F. D., King's College, London, and at Central Technical College, London. Jones, E. V., University, Birmingham. Kirkhope, T. B., Glasgow and West of Scotland Technical College. Munro, W. T., Glasgow and West of Scotland Technical College. Neville, H. A. D., B.Sc. (Lond.).

Peters, W. H., University College, Nottingham. Walker, J. H., Glasgow and West of Scotland Technical College. Wallis, T. E., B.Sc. (Lond.), Pharmaceutical Society's Laboratory. Wilson, C. J., Finsbury Technical College, London.

EXAMINATION IN GENERAL PRACTICAL CHEMISTRY FOR FELLOWSHIP.—Coppock, J. B., University College, Nottingham, and at Royal College of Science, London.

FINAL A.I.C. EXAMINATION.—Branch "A" (*Mineral Chemistry*): McLellan, B. G., Glasgow and West of Scotland Technical College. Northall-Laurie, D., King's College, London. O'Shaughnessey, F. R., A.R.C.Sc., Royal College of Science, London. Tainsh, P. W., Glasgow and West of Scotland Technical College. Taylor, T., Glasgow and West of Scotland Technical College. Scott Tebb, W., M.A., M.D. (Cantab.), Cambridge University and King's College, London. Wade, F., A.R.C.Sc., Royal College of Science, London. Watson, F. W., B.Sc. (Lond.), Glasgow and West of Scotland Technical College. Branch "B" (*Metallurgical Chemistry*): Abell, C. T., B.Sc. (Vict.), Owens College, Manchester. Comber, A. W., Finsbury Technical College, London. Branch "D" (*Organic Chemistry*): Dakin, H. D., B.Sc. (Vict.), Yorkshire College, Leeds. Hinks, E., King's College, London. Slator, A., B.Sc. (Lond.), University, Birmingham. Smith, R. E. B., B.Sc. (Lond.), University College, London. Branch "E" (*Analysis of Food and Drugs and of Water, including Examination in Therapeutics, Pharmacology and Microscopy*): Coysh, B. R., King's College, London. Ellis, A. W., University, Birmingham, and under J. K. Colwell, Esq., F.I.C. Jardin, D. S., A.R.C.Sc.I., Royal College of Science, Dublin. Macara, T., Glasgow and West of Scotland Technical College, and under Dr. J. Clark, F.I.C. Melling, S. E., Owens College, Manchester, and under A. H. Allen, Esq., F.I.C. Webster, J., University, Birmingham, and under J. K. Colwell, Esq., F.I.C. Branch "E" (*Analysis of Food and Drugs and of Water*): Brown, J. A., University College, Nottingham.

The examiners in Chemistry were Dr. Bernard Dyer, F.I.C., and Dr. W. Palmer Wynne, F.R.S., F.I.C. The examiner in Therapeutics, Pharmacology and Microscopy was Dr. A. P. Luff, F.R.C.P., F.I.C.

ARSENIC COMMITTEE.

THE Joint Committee of the Societies of Chemical Industry and Public Analysts have circulated a number of samples containing definite quantities of arsenic, and have carefully considered the results returned by the members. The Committee believe that these results justify them in prescribing a method for dealing with the various substances, but before taking this course they think it desirable to submit this method to somewhat more extended trials.